**ADDRESSING STUDENT MISCONCEPTIONS**

**ABOUT GASES AND ATMOSPHERIC CHEMISTRY**

1. Students beginning this unit may have the misconception that gases possess less mass (“lower weight”, or, “no weight”) than solids and liquids. When given a conceptual problem where, for example, a liquid sample of iodine with mass X in a closed container is allowed to evaporate, the students will state that the total mass of the system will decrease following evaporation.

One way to address this issue is to allow students to determine the mass of a two-litre soda bottle, and a sample (0.50 to 0.75 grams) of dry ice. The bottle is carefully sealed with tape after adding the dry ice, and the sample is allowed to sublime at room temperature. The students, then, determine the mass, after sublimation, and should observe that mass is conserved.

Safety Concerns:

The amount of dry ice added to the bottle should not exceed the recommended range, because otherwise the build-up of pressure from the gaseous carbon dioxide could cause the bottle to explode. Likewise, no other liquids or solids should be present in the bottle, for similar reasons.

2. Students tend to confuse average kinetic energy and molecular velocity. It is important for students to understand the concept that, when two gases of different molecular masses are at the same temperature, their average kinetic energies (E = 1/2 mv2) are the same, even though the average velocities of the molecules are different. Students may, also, be unable to differentiate the process of diffusion from the kinetic molecular theory that explains it.

An interesting demonstration for students, which illustrates the difference between molecular velocity and average kinetic energy, as determined by temperature, is the HCl/NH3 “molecular race”. Two clear plastic drinking straws are taped together, end-to-end; this will form the “track” for the race. Small samples of concentrated HCl and NH3 are placed in separate beakers; thermometers should indicate that they are both at the same temperature (this will, also, approximate equal temperatures for the two gases involved in the demonstration). A cotton swab is cut in half, and each cotton end is dipped in one of the two solutions, to a point of cotton saturation, but not dripping. Each cotton swab is placed in one end of the straw track, and the ends are sealed with tape. Students wait for the appearance of a ring of solid ammonium chloride, and measure how far from either end the rink appears. Students should find that the ring appears closer to the HCl end, reflecting the larger mass and, therefore, lower molecular velocity of the HCl. Students should be made aware of the difference in molecular mass of both compounds, using blackboard, or overhead, images of both molecules, or with model kits.

Safety Concerns:

Both aqueous solutions used are concentrated, and can irritate, or damage, skin and eyes. They should, therefore, be handled with caution, preferably, in a fume hood.

3. Students have many misconceptions related to kinetic molecular theory. A very common one concerns the behaviour of molecules of a substance when that substance is heated. Specifically, students sometimes think that molecules expand when a liquid evaporates to a gas, or when a gas is heated so that its volume increases.

There are several ways to address this. After introducing students to kinetic molecular theory, present them with the scenario where an air sample in an Erlenmeyer flask, which has a balloon stretched across the mouth, is heated so that the balloon begins to inflate (the teacher could actually perform this as a demonstration in front of the class). Ask the students to illustrate what they think happens to the molecules in the air to account for the balloon’s inflation. The students could, then, form groups and discuss their illustrations (think-pair-share). It would, then, be advisable to introduce a computer simulation to the students. Explore Learning has a “Gizmo” simulation called “Temperature and Particle Motion”, which allows them to observe the motion of molecules of an ideal gas at a range of temperatures (<http://www.explorelearning.com/index.cfm?method=cResource.dspView&ResourceID=555>).

If students do believe that gas molecules expand in size, the instructors can challenge them whether or not they think the masses of the molecules increase. If they believe this, the teacher can perform a demonstration comparing the mass of a flask partially filled with water, with a balloon stretched across the mouth to the same system after heating, so that some of the water has vaporized, and the balloon has become partially inflated with gaseous water molecules.

4. Following introduction to the ideal gas law, and the concept of molar volume, students may be tempted to think that one mole of any gas occupies 22.4 litres, under any circumstances, and use this figure, erroneously, as a shortcut to solving gas stoichiometry problems.

It is very important to stress to students that the value of 22.4 L/mol applies to ideal gases at standard temperature and pressure (STP). The teacher should explicitly lead the class through problems using the ideal gas law under a variety of ambient conditions. After every problem, the teacher should model what is the molar volume of the gas, beginning with an accurate representation of a volume of 22.4 L. This should be a cube of side length 28.2 cm, labelled “One mole”, “22.4 litres”, “**Only at STP**”. Careful preparation for in-class problem-solving should include at least some physical representation of the molar volume at conditions other than STP, using a variety of containers of appropriate size.

5. Some students commonly (intuitively) believe that air does not normally exert any pressure on objects because you can’t feel it. There are numerous ways to address this misconception. This could be handled in the very first lesson (after the “hook”). One method is to use the soda can crushing experiment, described in a previous section. Another way to demonstrate that air does, indeed, contain particles that can be compressed is to try to blow up a small balloon inside an empty two-litre soda bottle, both before and after cutting out the bottom of the vessel. Students should understand why their ability to inflate the balloon inside the closed system is diminished, because of limitations in compressing the air already present in the bottle.

6. Chemistry students often have difficulty with conceptual understanding of gas laws, and how they are applied in real-world situations. When given quantitative problems to solve, some automatically try to plug the given information into one of the gas law equations. It is important to try to reason what is happening at the molecular level, using kinetic molecular theory. Graphic organizers can be helpful in this regard; for instance, blank figures of a piston provide the opportunity to graphically represent a Boyle’s Law application, where students can make the alterations to the volume of the piston, and draw the gas molecules accordingly exerting more or less pressure.